Chapter-7

Lesson-1

I – The Periodic Table

- In 1803 John Dalton introduced the modern atomic theory that elements were composed of atoms and that the atoms were different for different elements. In 1808, Dalton created a system for representing the different elements so that compounds could be more easily written.
  - Old alchemy symbols:
    copper, ♀ ; antimony, ♁ ; iron, ♂ ; silver, ☽ ; gold, ◦ ; tin, ♃
  - Dalton's symbols:
    copper Ⓒ , carbon ● , oxygen ○ , hydrogen ⊙ , water ⊙○

- In 1811, Jöns Jakob Berzelius introduced the modern system of chemical symbols for the elements. Each element was assigned a symbol of one or two letters, the first always being capitalized. Berzelius' system would undergo some refining over the years. For example H^2O became H_2O. In 1828, Berzelius published the first accurate list of atomic masses for the elements.

- In 1829, Wolfgang Dobereiner developed a system of periodically returning characteristics of the elements, a precursor of the periodic table. Dobereiner also discovered a number of sets of three elements (he called them "triads") in which the members have similar properties. Furthermore, the average of the masses of the lightest and heaviest members is close to the mass of the other member! Four such triads are: Li, Na, & K; Ca, Sr, & Ba; Cl, Br, & I; and S, Se, & Te.

- In 1863, John Newlands announced what he called the "Law of Octaves" which stated: when the elements are arranged in ascending order by mass, every eighth element has similar properties. In 1864, he publishes a periodic table (see next page).

Newlands (1864)						
						H
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca	***Cr***	Ti	***Mn***	***Fe***	***Co, Ni***
Cu	Zn	***Y***	***In***	As	Se	Br
Rb	Sr	***La, Ce***	Zr	Nb, ***Mo***	***Ru, Rh***	***Pd***
Ag	Cd	***U***	Sn	Sb	Te	I
Cs	Ba, ***V***					

Elements in boldface are out of place.

- In 1869 Dmitri Mendeleev published a table of the elements. In it the elements were listed, in ascending order by mass, in columns and in rows. Mendeleev further claimed that the elements in the same column had similar chemical and physical properties. It was this fact that led Mendeleev to construct his table in the first place. This also led him to state his **periodic law**: when the elements are placed in ascending order by mass, at set intervals the general properties of the elements repeat (i.e. the properties **periodically** repeat themselves)!

Mendeleev (as revised 1871)							
I	II	III	IV	V	VI	VII	VIII
R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7	RO_4
H							
Li	Be	B	C	N	O	F	
Na	Mg	Al	Si	P	S	Cl	
K	Ca	*-	Ti	V	Cr	Mn	Fe,Co,Ni
Cu	Zn	*-	*-	As	Se	Br	Ru,Rh,Pd
Ag	Cd	In	Sn	Sb	Te	I	
Cs	Ba						

***Undiscovered elements whose properties Mendeleev predicted.**

- Mendeleev's table was not perfect. **First**, several of the elements had to be placed out of order by mass so that they would fall in the column with other elements of similar properties. Mendeleev could not explain why this happened. **Second**, Mendeleev's table had some empty spots because no known element fit the column and row based on its properties. Mendeleev claimed that eventually elements would be discovered to fill these "holes".

- Mendeleev predicted three yet-to-be-discovered elements' chemical and physical properties. At first, the scientific community ridiculed his table as nonsense. However, after all three elements were discovered one by one that fit Mendeleev's predictions, the world began to take notice of his periodic table.

- By 1885 Mendeleev's Table was widely accepted. In bears noting that, in 1870, Julius Lothar Meyer independently published a table very similar to Mendeleev's Table. However Meyer did not go on to predict the properties of undiscovered elements and thus Mendeleev is given the credit for the modern periodic table.

- It would take over forty years for Mendeleev's first problem to be solved. After the proton was discovered by Ernest Rutherford (1914) and their exact number in each element determined by Henry Gwyn-Jeffreys Moseley (1914), it was revealed that Mendeleev's table was in fact perfect! Mendeleev's mistake was in ordering the elements by their atomic mass and not by their atomic number (number of protons). Of course he would never know this because he died in 1907.

- Today, the elements in the modern periodic table are placed in ascending order of their the atomic number (number of protons). **The modern periodic law states**: <u>the properties of the elements will periodically repeat themselves when they are placed in ascending order by atomic number</u>.

II –Sizes of Atoms and Ions

A) Effective Nuclear Charge

- The ***effective nuclear charge*** (**Z_{eff}**) of any valence electron can be determined by the expression

$$\mathbf{Z_{eff} = Z - S}$$

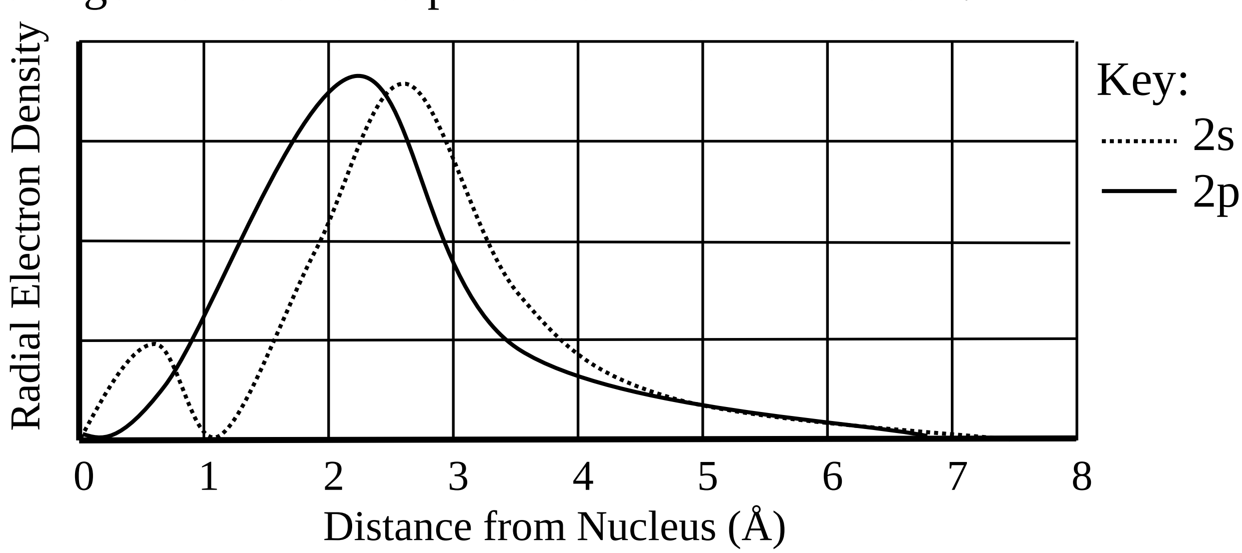
where **Z** is the actual nuclear charge (atomic number) and **S** is a positive number called a screening constant.

- The ***effective nuclear charge*** (**Z_{eff}**) of a valence electron can be thought of as ***how much a valence electron "sees" the nucleus***.

Actual Effective Nuclear Charge (Z_{eff}) on Valence s-Electrons (1st value) and on Valence p-Electrons (2nd value)

H (-)	He (1.69)					H (-)	He (1.69)
Li (1.28)	Be (1.91)	B (2.58) (2.42)	C (3.22) (3.14)	N (3.85) (3.83)	O (4.49) (4.45)	F (5.13) (5.10)	Ne (5.76) (5.76)
Na (2.51)	Mg (3.31)	Al (4.12) (4.07)	Si (4.90) (4.29)	P (5.64) (4.89)	S (6.37) (5.48)	Cl (7.07) (6.12)	Ar (7.76) (6.76)
K (3.50)	Ca (4.40)	Ga (7.07) (6.22)	Ge (8.04) (6.78)	As (8.94) (7.45)	Se (9.76) (8.29)	Br (10.55) (9.03)	Kr (11.32) (9.77)
Rb (4.98)	Sr (6.07)	In (9.51) (8.47)	Sn (10.63) (9.10)	Sb (11.61) (9.99)	Te (12.54) (10.81)	I (13.40) (11.61)	Xe (14.22) (12.42)

- A valence s-electron in an atom can "see" more of the nuclear charge than a valence p-electron in the same atom.



- This is because s-orbital electrons in a many electron atom have a noticeable probability of being up close to the nucleus whereas p-orbital electrons do not. Thus s-orbital electrons are less effectively screened by the core electrons than the p-orbital electrons in the same PEL. S-orbital electrons “see” more of the nuclear charge than p-orbital electrons in the same PEL.
- This also explains why the s-orbital is lower in energy than the p-orbitals for the same PEL. If the s-orbital electrons have a larger Z_{eff} than the p-orbital electrons then the greater attraction between the s-orbital electrons and the nucleus leads to lower energy (***s-orbital electrons also shield p-orbital electrons***).

B) Atomic Radius

- In general, as one moves from left to right across the Periodic Table, the atomic radius of the elements decreases.
- When one moves down a group on the Periodic Table, the atomic radius increase.
- These trends can be explained in terms of the number of ***PEL’s*** (**n**) in an atom and of the ***effective nuclear charge*** (**Z_{eff}**) acting on the valence electrons in an atom.
- ***Ideally, the value of S is equal to the number of core (kernel) electrons in an atom because the core electrons would (ideally) completely shield the valence electrons from the nucleus.***

Ideal Effective Nuclear Charge (Z_{eff}) on Valence Electron(s)

H (1) 37	He (2) 32	***And Atomic Radius (pm)***				H (1) 37	He (2) 32
Li (1) 134	Be (2) 90	B (3) 82	C (4) 77	N (5) 75	O (6) 73	F (7) 71	Ne (8) 69
Na (1) 154	Mg (2) 130	Al (3) 118	Si (4) 111	P (5) 106	S (6) 102	Cl (7) 99	Ar (8) 97
K (1) 196	Ca (2) 174	Ga (3) 126	Ge (4) 122	As (5) 119	Se (6) 116	Br (7) 114	Kr (8) 110
Rb (1) 211	Sr (2) 192	In (3) 144	Sn (4) 141	Sb (5) 138	Te (6) 135	I (7) 133	Xe (8) 130

- **Within in Period**: The atomic radius decreases across a period due to an increase in the effective nuclear charge acting on the valence electrons. The increase in the effective nuclear charge steadily draws the valence electrons closer to the nucleus, causing the radius to decrease.
- **Within a Group**: The atomic radius increases down a group due to an increase in the number of PEL's. The valence electrons have a greater and greater probability of being farther from the nucleus, causing the radius to increase.

C) Ionic Radius

- Cations are smaller than their neutral atoms. This is primarily due to the loss of all the electrons in the outermost PEL of the atom. The remaining electrons are closer to the nucleus, causing the radius to decrease.
- Anions are larger than their neutral atoms. This primarily due to electron-electron repulsions causing the valence electrons to spread out and the radius to increase.
- For ions in the same group and carrying the same charge, the radius increases as the number of PEL's increase.
- For ***isoelectronic ions*** (having the same electron configuration), the radius decreases as the atomic number **(# of protons)** increases.

Q1: Arrange the following atoms in order of increasing atomic radius: Al, Ca, and Ga.

A1: Al < Ga< Ca

Q2: Arrange the following species in order of increasing atomic radius: K, K^{1+}, Na^{1+}

A2: $Na^{1+} < K^{1+} < K$

Q3: Arrange the following ions in order of increasing atomic radius: S^{2-}, Cl^{1-}, K^{1+}, Ca^{2+}, Sc^{3+}

A3: $Sc^{3+} < Ca^{2+} < K^{1+} < Cl^{1-} < S^{2-}$

Chapter-7

Lesson-2

III – Ionization Energy

- ***Ionization energy*** (IE) is the minimum amount of energy that an atom or ion must absorb to remove an electron.
- The ***1st ionization energy*** is the amount of energy that a neutral atom must absorb to convert it to a 1+ ion.

$$\mathbf{Be^{0}_{(g)} + 899\ kJ \rightarrow Be^{1+}_{(g)} + 1e^{-}}$$

1st Ionization Energies of the Main-Group Elements (kJ/mol)

H 1312							He 2372
Li 520	Be 899	B 801	C 1086	N 1402	O 1314	F 1681	Ne 2081
Na 496	Mg 738	Al 578	Si 786	P 1012	S 1000	Cl 1251	Ar 1521
K 419	Ca 590	Ga 579	Ge 762	As 947	Se 941	Br 1140	Kr 1351
Rb 403	Sr 549	In 558	Sn 709	Sb 834	Te 869	I 1008	Xe 1170
Cs 376	Ba 503	Tl 589	Pb 718	Bi 703	Po 812	At —	Rn 1037

- The ***2nd ionization energy*** is the amount of energy that an atom or ion must absorb to convert it to a 2+ ion.

$$\mathbf{Be^{1+}_{(g)} + 1760\ kJ \rightarrow Be^{2+}_{(g)} + 1e^{-}}$$

- Ionization energies for an element increase in magnitude as successive electrons are removed. This trend arises because with each successive removal, an electron is being pulled away from an increasingly more positive ion with a greater effective nuclear charge, requiring increasingly more energy.

- The ***3rd ionization energy*** is the amount of energy that an atom or ion must absorb to convert it to a 3+ ion.

$$\mathbf{Be^{2+}_{(g)} + 14850\ kJ \rightarrow Be^{3+}_{(g)} + 1e^-}$$

- There is a sharp rise in the ionization energy when the first kernel electron is removed. This occurs because kernel electrons are closer to the nucleus and have a greater Z_{eff} value.

Successive Ionization Energies for Period-2

Li	Be	B	C	N	O	F	Ne
520	*899*	*801*	*1086*	*1402*	*1314*	*1681*	*2081*
7300	*1760*	*2430*	*2350*	*2860*	*3390*	*3370*	*3950*
	14850	*3660*	*4620*	*4580*	*5300*	*6050*	*6120*
		25020	*6220*	*7480*	*7470*	*8410*	*9370*
			37830	*9440*	*10980*	*11020*	*12180*
				53270	*13330*	*15160*	*15240*
					71330	*17870*	*20000*
						92040	*23070*
							115380

- When an electron is removed from an atom or ion, it is removed from the orbital having the largest principal quantum number.

Sc^0 ($1s^22s^22p^63s^23p^64s^23d^1$) $\Rightarrow$ Sc^{1+} ($1s^22s^22p^63s^23p^64s^13d^1$)

Sc^{1+} ($1s^22s^22p^63s^23p^64s^13d^1$) $\Rightarrow$ Sc^{2+} ($1s^22s^22p^63s^23p^63d^1$)

Sc^{2+} ($1s^22s^22p^63s^23p^63d^1$) $\Rightarrow$ Sc^{3+} ($1s^22s^22p^63s^23p^6$) ***Stable octet***

Sc^{3+} ($1s^22s^22p^63s^23p^6$) $\Rightarrow$ Sc^{4+} ($1s^22s^22p^63s^23p^5$) ***Yikes!***

Q1: What is the abbreviated electron configuration for Cu^{2+}?

A1: [Ar]$4s^03d^9$ or [Ar]$3d^9$

Q2: What is the abbreviated electron configuration for Fe^{3+}?

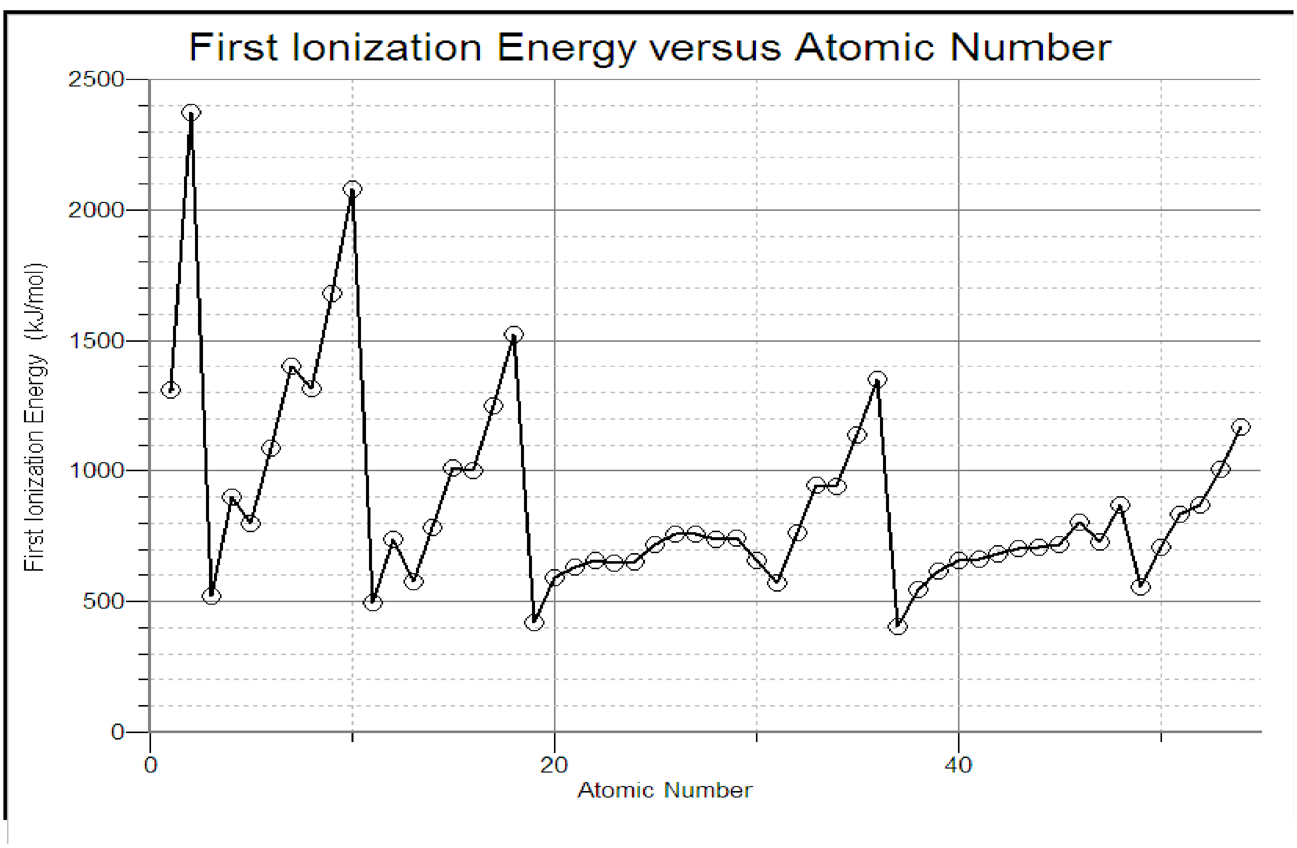
A2: [Ar]$4s^03d^5$ or [Ar]$3d^5$

Q3: What is the abbreviated electron configuration for I^{1-}

A3: [Kr]$5s^24d^{10}5p^6$

- Within each group on the Periodic Table, the ionization energy generally decreases. This is primarily due to the addition of a new PEL in which the valence electrons are farther from the nucleus.
- Within each period on the Periodic Table, the ionization energy generally increases. This is due to an increase in the effective nuclear charge that the valence electrons experience as well as the decrease of the atomic radius.
- The pronounced irregularities seen in period-2 and period-3 can be easily explained.

- The ionization energy for a ***boron atom*** ([He]$2s^2 2p^1$) is lower than the ionization energy for a ***beryllium atom*** ([He]$2s^2$) because the 2p sublevel is at a higher energy than the 2s. Thus it requires less energy to remove the 2p electron from boron than a 2s electron from beryllium even though boron's radius is smaller and its effective nuclear charge is larger than beryllium's.
- The ionization energy for an ***oxygen atom*** ([He]$2s^2 2p^4$) is lower than the ionization energy for a ***nitrogen atom*** ([He]$2s^2 2p^3$) because of the repulsion that exists between the paired electrons in the 2p-orbital of oxygen makes it easier to remove an electron.



IV – Electron Affinity

- ***Electron affinity*** (EA) is the change in energy when 1 mole of gaseous atoms gains a 1 mole of electrons.
- ***Electron Affinity*** measures the attraction (or affinity) of the atom for added electrons.

H -73	***1st Electron Affinity (kJ/mol)***						He > 0
Li -60	Be > 0	B -27	C -122	N > 0	O -141	F -328	Ne > 0
Na -53	Mg > 0	Al -43	Si -134	P -72	S -200	Cl -349	Ar > 0
K -48	Ca - 2 (≈0)	Ga -30	Ge -119	As -78	Se -195	Br -325	Kr > 0
Rb -47	Sr -5 (≈0)	In -30	Sn -107	Sb -103	Te -190	I -295	Xe > 0

- A negative electron affinity indicates that energy is released as the anion is formed and that the anion is at a lower energy.
- The greater the attraction between a given atom and an added electron, the more negative the atom's affinity will be.
- ***Halogens*** have the most negative electron affinity. By gaining an electron, a halogen forms the stable octet configuration of a noble gas.
- **EXCEPTIONS:** ***Noble gases***, smaller ***alkaline earth metals***, and ***nitrogen*** have positive electron affinities. This indicates that these atoms would become less stable as anions.
- In the case of the noble gases, the addition of an electron would require the electron to reside in a higher PEL orbital.
- For the alkaline earth metals, the addition of an electron would require the electron to reside in a higher energy orbital.
- When a group 15 family member gains an electron, a half-filled p-orbital becomes filled and destabilizing electron-electron repulsions result. Since nitrogen is a small atom, the electron-electron repulsions are greater and more influencing than with its larger family members.

Chapter-7

Lesson-3

V – Properties of Metals, Nonmetals, and Metalloids

A) Metallic Properties (Character):

- ***Metallic character*** is the extent to which an element has properties typical of a metal. A metal tends to …
  - be a shade of gray
  - be solid at room temperature
  - be shiny (lustrous)
  - be malleable and ductile
  - be a good conductor of heat and electricity
  - lose electrons easily during reactions

- Highly active ***metals*** react with water to form hydroxide bases and hydrogen gas [***Li, Rb, K, Cs, Ba, Sr, Ca, & Na***].

  $2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$

- Most ***metals*** react with acids to form a salt and hydrogen gas.

  $2\ K + 2\ HCl \rightarrow 2\ KCl + H_2$

- ***Metals*** react with oxygen to form metallic oxides.

  $2\ Mg + O_2 \rightarrow 2\ MgO$

- ***Metallic oxides*** react with acids to form a salt and water.

  $CuO + HC_2H_3O_2 \rightarrow Cu(C_2H_3O_2)_2 + H_2O$

- ***Metallic oxides*** react with water to form hydroxide bases.

  $Li_2O + H_2O \rightarrow 2\ LiOH$

- ***Metallic character*** increases down a group but decreases across (left to right) a period. ***Francium is the "most" metallic of the metals***.

- In ***metallic oxides***, the higher the metallic character of the metal, the more basic its aqueous mixture will be.

B) Nonmetallic Properties (Character):

- ***Nonmetallic character*** is the extent to which a substance has properties typical of a nonmetal. A nonmetal tends to …
    - be colorful
    - be gaseous at room temperature
    - be dull (not lustrous) when solid
    - be brittle when solid
    - be a poor conductor of heat and electricity
    - gain electrons easily during reactions

- ***Nonmetals*** react with metals to form ionic compounds.

  $2\ Cu + S \rightarrow Cu_2S$ or $Na + Cl \rightarrow NaCl$

- ***Nonmetals*** react with other nonmetals to form molecular compounds.

  $C + 2\ H_2 \rightarrow CH_4$ or $2\ P + 5\ Cl_2 \rightarrow 2\ PCl_5$

- ***Nonmetals*** react with oxygen to form nonmetallic oxides.

  $S + O_2 \rightarrow SO_2$

  $2\ N_2 + 5\ O_2 \rightarrow 2\ N_2O_5$

- ***Nonmetallic oxides*** react with water to form acids.

  $H_2O + SO_2 \rightarrow H_2SO_3$

  $H_2O + N_2O_5 \rightarrow 2\ HNO_3$

- ***Nonmetallic oxides*** react with bases to form a salt and water.

  $CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$

  $SO_3 + 2\ NaOH \rightarrow Na_2SO_4 + H_2O$

- ***Nonmetallic character*** decreases down a group but increases across a period. ***Fluorine is the "most" nonmetallic of the nonmetals***.

- In ***nonmetallic oxides***, the higher the nonmetallic character of the nonmetal, the more acidic its aqueous mixture will be.

- ***Metalloids*** are elements that exhibit partial nonmetallic character. [***B, Si, Ge, As, Sb, & Te*** (and At?)]

VI – Group Trends in Metals and Nonmetals

A) The Alkali and Alkaline Earth Metals (Groups 1 & 2):

1) Physical Properties:

- Hardness:
  - ***Alkali metals*** are soft while ***alkaline earth metals*** are harder.
- Melting Point:
  - ***Alkali metals*** have low m.p. (< 400^{o}C).
    The m.p. decreases down the group.
  - ***Alkaline earth metals*** have medium m.p. (400^{o}C to 1000^{o}C) except for Be which is high.
    The m.p. generally decrease down the group.
- Ionization Energy:
  - ***Alkali metals and alkaline earth metals*** have low ionization energies.
- Atomic Radius:
  - In their period, ***alkali metals and alkaline earth metals*** have large atomic radii.
- Density:
  - ***Alkaline earth metals*** have a greater density than their neighboring ***alkali metals*** but both have relatively low densities as metals.
- Specific Heat:
  - ***Alkali metals and alkaline earth metals*** have relatively large specific heat values that decrease both down the group and across the period.
- Flame Test Color:
  - Li is red, Na is yellow-orange, K is violet, Ca is red-orange, Sr is crimson, and Ba is yellow-green. ***Fireworks and road flares***

2) Chemical Properties:

- Reactions with Water:
  - ***Alkali metals*** react vigorously with water to form a base and hydrogen gas.

    $2\ Na_{(s)} + 2\ H_2O \rightarrow 2\ NaOH_{(s)} + H_2$

  - Only the heavier ***Alkaline earth metals*** (Ca, Sr, & Ba) react vigorously with water to form a base and hydrogen gas.

    $Ca_{(s)} + 2\ H_2O \rightarrow Ca(OH)_{2(s)} + H_2$

    $Ba_{(s)} + 2\ H_2O \rightarrow Ba(OH)_{2(s)} + H_2$

  - ***Mg*** reacts slowly in water but reacts readily with steam*.

    $^*Mg_{(s)} + 2\ H_2O_{(g)} \rightarrow Mg(OH)_{2(s)} + H_2$

- Reactions with Oxygen:
  - ***Normal Oxides***: $[O^{2-}]$

    Lithium reacts with oxygen to give the normal oxide (M_2O).

    $4\ Li_{(s)} + O_{2(g)} \rightarrow 2\ Li_2O_{(s)}$

    Group-2 metals react with oxygen to give the normal oxide (MO).

    $2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$ (at very high temperatures)

    $2\ Sr_{(s)} + O_{2(g)} \rightarrow 2\ SrO_{(s)}$ (at high temperatures)

    $2\ Ba_{(s)} + O_{2(g)} \rightarrow 2\ BaO_{(s)}$

  - ***Peroxides***: $[O_2^{2-}]$

    Sodium reacts directly with oxygen to form the peroxide.

    $2\ Na_{(s)} + O_{2(g)} \rightarrow Na_2O_{2(s)}$

  - ***Superoxides***: $[O_2^{1-}]$

    Potassium, ***rubidium***, and ***cesium*** form the superoxide when directly reacting with oxygen.

    $K_{(s)} + O_{2(g)} \rightarrow KO_{2(s)}$

    $Rb_{(s)} + O_{2(g)} \rightarrow RbO_{2(s)}$

    $Cs_{(s)} + O_{2(g)} \rightarrow CsO_{2(s)}$

B) The Transition Metals (Group 3 to Group 12):

1) Physical Properties:

- Atomic Radius:
  - Across a period, the radii of the first transition metals generally decrease and then the radii of the last two to four transition metals generally increase slightly.
- Density:
  - Across a period, the densities of the transition metals generally increase and then the densities of the last two to four transition metals generally decrease slightly.
- Melting Point:
  - Across a period, the melting points of the transition metals generally increase then the last two to four decrease.
- Ionization Energy:
  - Across a period, the ionization energies of the transition metals generally increase.

2) Chemical Properties:

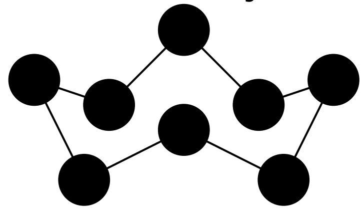
- Relative Reactivity:
  - The transition metals are usually less reactive than alkali or alkaline earth metals.
- Multiple Compounds:
  - The transition metals usually form more than one cation and thus more than one compound with the same element.
- Colored Compounds:
  - The transition metals form compounds with a wide variety of color. ***Exceptions are the compounds of scandium (Sc^{3+}), titanium IV (Ti^{4+}), and zinc (Zn^{2+}), which are always white.***

Na_2CrO_4 - yellow.	$FeSO_4$ - light green.	$CoCl_2 \bullet 6H_2O$ - pink.
$Na_2Cr_2O_7$ - orange.	$Fe(OH)_3$ - red.	$CuSO_4 \bullet 5H_2O$ - blue.
$KMnO_4$ - purple.	$FeSCN^{2+}$ - blood red.	$NiCl_2 \bullet 6H_2O$ - green.

C) The Chalcogens and Halogens (Groups 16 & 17):

1) Physical Properties:

- The ***chalcogens*** include 3 nonmetals, a metalloid, and a metal:
  - oxygen, sulfur, selenium, tellurium, and polonium
- The ***halogens*** include 4 nonmetals and 1 nonmetal/metalloid:
  - fluorine, chlorine, bromine, iodine, and astatine
- Melting Point:
  - ***Chalcogens*** have medium to low m.p. ($< 600^oC$). The m.p. increase down the group (except for Po which is less than Te).
  - ***Halogens*** have medium to low m.p. ($< 600^oC$). The m.p. increases down the group.
- Ionization Energy:
  - ***Chalcogens*** and ***halogens*** have high ionization energies.
  - Down their group, the ionization energies decrease.
  - In a period, the ***halogen*** has the higher ionization energy.
- Atomic Radius:
  - For their period, ***chalcogens*** and ***halogens*** have small atomic radii with the halogen bring smaller than the chalcogen.
- Elemental State:
  - Pure sulfur is a bright yellow solid that exists as S_8 molecules. The S_8 molecule is an eight-membered puckered ring. However, pure sulfur is usually written as $S_{(s)}$.



  - Pure oxygen occurs in two allotropes: O_2 and O_3 (ozone).
  - Pure chlorine is a greenish-yellow gas (Cl_2).
  - Pure bromine is a red liquid (Br_2).
  - Pure iodine is a dark purple solid (I_2).

2) Chemical Properties:

- Pure ***chalcogens*** and ***halogens*** are highly reactive. In a period, the ***halogens*** are more reactive than the ***chalcogen***.

- Preparation of Oxygen

  Catalyzed decomposition of hydrogen peroxide

$$2\ H_2O_{2(aq)} \xrightarrow{\textbf{catalyst}} 2\ H_2O_{(\ell)} + O_{2(g)}$$

  Thermal decomposition of potassium chlorate

$$2\ KClO_{3(s)} \xrightarrow{\textbf{heat}} 2\ KCl_{(s)} + 3\ O_{2(g)}$$

- Reactions Involving Ozone

  Formation of Ozone during Thunder Storms

$$3\ O_{2(g)} \rightarrow 2\ O_{3(g)}$$

  Catalyzed Photo-decomposition of Ozone

$$2\ O_{3(g)} \rightarrow 3\ O_{2(g)}$$

- Reaction of a Halogen by a More Reactive Halogen

$$Br_2 + 2\ I^- \rightarrow I_2 + 2\ Br^-$$

$$Cl_2 + 2\ Br^- \rightarrow Br_2 + 2\ Cl^- \qquad [F_2 > Cl_2 > Br_2 > I_2]$$

$$F_2 + 2\ Cl^- \rightarrow Cl_2 + 2\ F^-$$

- Preparation of Chlorine

$$MnO_{2(s)} + 4\ HCl_{(aq)} \rightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2\ H_2O_{(\ell)}$$

- Reaction of Fluorine and Chlorine in Water

$$2\ F_2 + 2\ H_2O \rightarrow 4\ HF + O_2$$

$$Cl_2 + H_2O \rightarrow HCl + HOCl\ (HClO)$$

- Reaction of Halogens and Sulfur with (most) Metals

$$2\ Na + Cl_2 \rightarrow 2\ NaCl$$

$$2\ Al + 3\ Br_2 \rightarrow 2\ AlBr_3$$

$$Ca + F_2 \rightarrow CaF_2$$

$$2\ Cu + S \rightarrow Cu_2S$$

D) The Noble Gases (Group 18):

- Until the early 1960's, the group 18 elements were called the inert gases because in nature these elements are extremely unreactive.
- All Noble gases have a filled valence and exist as monatomic molecules.
- Only the heaviest of the noble gases (Ar, Kr, and Xe) form compounds.

1) Physical Properties:

- Ionization Energy
  - Noble gases have large ionization energies which decrease down the group.
- Atomic Radius
  - Noble gases have small radii which increase down the group.
- Density
  - Noble gases have low densities (at 1 atm and 298 K) which increase down the group.
- Boiling Point
  - Noble gases have low boiling points which increase down the group.

2) Chemical Properties:

- Reaction of Xenon with Fluorine, Chlorine, and Oxygen

$$Xe + 2\,F_2 \rightarrow XeF_4$$
$$Xe + Cl_2 \rightarrow XeCl_2$$
$$Xe + 2\,F_2 + O_2 \rightarrow XeO_2F_4$$

- Reaction of Krypton with Fluorine

$Kr + F_2 \rightarrow KrF_2$ (*the only known compound*)

- Reaction of Argon

$H_2 + 2\,Ar + F_2 \rightarrow 2\,HArF$ (*the only known compound*)

- *Radon is so radioactive so only limited study has taken place*.